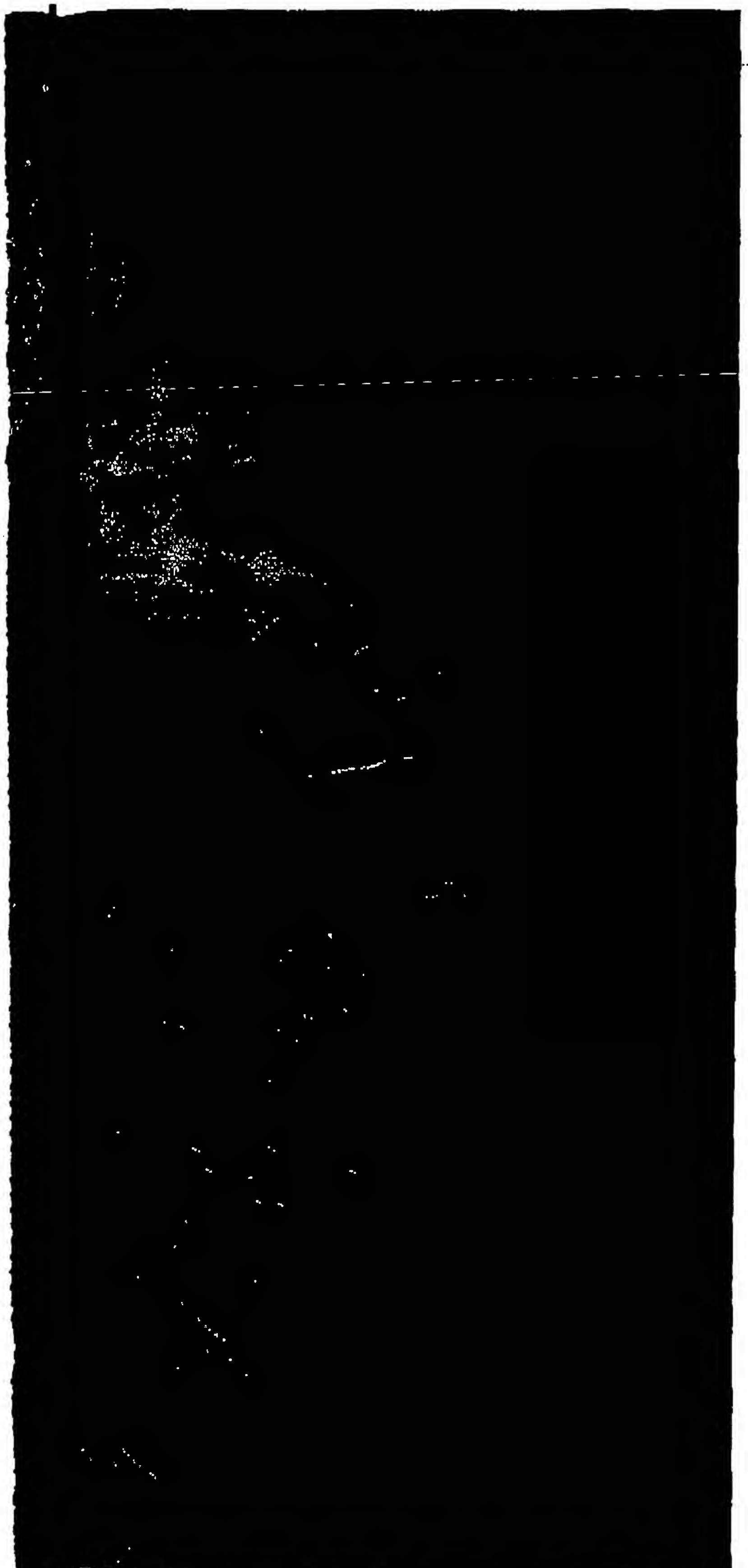

Chemistry for Engineers and Scientists

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(NH_4Cl) is added to sufficient water to make up one liter of solution.
Answer: $[H_3O^+] = 1.3 \times 10^{-5}$; pH = 4.89; $[OH^-] = 7.7 \times 10^{-10}$. ■

Buffers

When an acid or a base, even in relatively small amounts, is added to water, there is a large change in the pH of the solution. For example, addition of as litle as 0.010 mol of HCl lowers the pH of solvent water from 7 to 2, a change of 5 pH units. A **buffer** is a solution that is capable of absorbing excess acid or base without major changes in pH. Because such a solution must be able to react with a large fraction of added acid or base, the buffer solution must contain both an acid and a base. However, neither can be strong, since they would simply neutralize each other:

$$H_3O^+ \text{ (aq)} + OH^- \text{ (aq)} \longrightarrow 2\ H_2O \text{ (liq)}$$

Instead, a buffer must contain a weak acid and a weak base. More specifically, a buffer must contain a weak acid and its conjugate base.

For example, a liter of solution containing 0.50 mol of formic acid (HCOOH) and 0.50 mol of sodium formate (HCOONa) can provide effective buffering action. Before any dissociation occurs, the resulting solution is 0.50 M in formic acid and sodium formate. To determine the $[H_3O^+]$ and pH of this solution, keep in mind that sodium formate is a freely soluble salt, which dissociates essentially completely:

$$HCOONa \text{ (s)} \longrightarrow Na^+ \text{ (aq)} + HCOO^- \text{ (aq)}$$

Thus, the solution is 0.50 M in formate ion. HCOOH ionizes according to the following equation:

$$HCOOH \text{ (aq)} + H_2O \text{ (aq)} \rightleftharpoons H_3O^+ \text{ (aq)} + HCOO^- \text{ (aq)}$$

$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]} = 1.8 \times 10^{-4}$$

$$[H_3O^+] = (1.8 \times 10^{-4}) \frac{[HCOOH]}{[HCOO^-]}$$

If we let $[H_3O^+]$ formed as a result of this ionization be x, then, assuming $x \ll 0.500$

$$x = [H_3O^+] = (1.8 \times 10^{-4}) \left(\frac{0.500}{0.500}\right) = 1.8 \times 10^{-4}\ M$$

$$pH = -\log [H_3O^+] = -\log (1.8 \times 10^{-4}) = 3.74$$

Now suppose we dilute this buffer solution by adding enough water to make 2.0 L. The concentrations of HCOOH and $HCOO^-$ drop accordingly, to 0.25 M, and we may write the equilibrium expression as follows:

$$[H_3O^+] = (1.8 \times 10^{-4}) \left(\frac{0.25}{0.25}\right) = 1.8 \times 10^{-4}\ M$$

Note that the $[H_3O^+]$ and pH are unchanged, demonstrating one very important feature of buffer solutions: moderate dilution of a buffer solution does not significantly change its pH.

Let's now go back to our original, undiluted 1.0 L of buffer solution and add 0.010 mol of strong acid. What effect does this have on the $[H_3O^+]$ and pH of the resulting solution? Remember what happened when we